Acid-Base Titration

Pre-Lab Discussion

In the chemistry laboratory, it is sometimes necessary to experimentally determine the concentration of an acid solution or a base solution. A procedure for making this kind of determination is called an **acid-base titration.** In this procedure, a solution of known concentration, called the *standard solution,* is used to neutralize a precisely measured volume of the solution of unknown concentration to which one or two drops of an appropriate acid-base indicator have been added. If the solution of unknown concentration is acidic, a standard base solution is added to the acid solution until it is neutralized. If the solution of unknown concentration is basic, a standard acid solution is added to the base solution until it is neutralized.

When carrying out an acid-base titration, you must be able to recognize when to stop adding the standard solution, that is, when neutralization is reached. This is the purpose of the acid-base indicator mentioned above. A sudden change in color of the indicator signals that neutralization has occurred. At this point, the number of hydronium ions from the acid is equal to the number of hydroxide ions from the base. The point at which this occurs is called the *end point* of the titration. When the end point is reached, the volume of the standard solution used is carefully determined. Then, the measured volumes of the two solutions and the known concentration of the standard solution can be used to calculate the concentration of the other solution. The following steps tell how to calculate the unknown concentration:

1. Write the balanced equation for the reaction. From the coefficients, determine how many moles of acid reacts with 1 mole of base (or vice versa). Use the coefficients to form a mole ratio.
2. *If the mole ratio is 1:1,* the following relationship can be used to calculate the unknown concentration:

$$M_a \times V_a = M_b \times V_b$$

where M_a = molarity of the acid solution
M_b = molarity of the base solution
V_a = volume of the acid solution
V_b = volume of the base solution

The equation for this relationship can be rewritten to find the solution of unknown concentration. For example, if the molarity of the base were unknown, the equation would be

$$M_b = \frac{M_a \times V_a}{V_b}$$

3. *If the mole ratio is not 1:1,* the calculation of the unknown molarity is slightly more complicated. For example, if 2 moles of standard acid solution is needed to neutralize 1 mole of base of unknown concentration, the following relationship exists:

$$M_a \times V_a = 2(M_b \times V_b)$$

The 2 in this equation is known as the *mole factor.*

In Part A of this experiment, you will determine the molarity of a solution of NaOH by titrating it with a standard solution of HCl. The equation for this reaction is

$$HCl + NaOH \rightarrow NaCl + H_2O$$

Because the mole relationship of H^+ to OH^- is 1:1, no mole factor will be needed in your calculations.

In Part B of the experiment, you will titrate household white vinegar. Most commercial vinegars contain at least 4% acetic acid by weight. You will use the NaOH solution whose molarity you determined in Part A for the titration of the vinegar.

Up to this point in your laboratory work, most of your quantitative experiments have required you to calculate mass relationships. This is known as *gravimetric analysis.* Titration requires you to use volume relationships, a technique known as *volumetric analysis.*

This experiment should lead to a better understanding of the properties of acids and bases, neutralization reactions, and titration techniques.

Purpose

HF

Determine the molarity of a NaOH solution by titrating it with a standard HCl solution. Determine the molarity of a sample of white vinegar.

Equipment

burets, 50-mL (2)
buret stand
double buret clamp
graduated cylinder, 10-mL
Erlenmeyer flask, 250-mL
beakers, 250-mL (2)
dropper pipet
pipet, 10-mL
suction bulb
safety goggles
lab apron or coat

Materials

0.100 *M* HCl (standard solution)
NaOH (concentration unknown)
phenolphthalein
distilled water
detergent solution
white vinegar

Safety

Follow all precautions for working with acids and bases. Note the caution alert symbols here and with certain steps in the "Procedure." Refer to page xi to review the precautions associated with each symbol. Always wear safety goggles and a lab coat or apron when working in the lab.

Procedure

PART A TITRATION OF BASE OF UNKNOWN CONCENTRATION

1. Wash two burets with detergent solution. Rinse them thoroughly, first with tap water, then with distilled water.
2. Obtain about 100 mL of standard acid solution in a clean, dry 250-mL beaker. Obtain about the same amount of the base of unknown concentration in a second 250-mL beaker. **CAUTION:** *Handle these solutions with care. They can cause painful burns if they come in contact with the skin.*
3. Pour about 10 mL of acid into one buret and rinse the inside surface of the buret thoroughly. Allow the acid to run out the buret tip. Fill the buret to slightly above the 0.0-mL mark with acid. Then allow the acid to flow out the buret tip until the bottom of the meniscus is at the 0.0-mL mark (see Figure 39-1). Be sure there are no bubbles in the tip. If bubbles are present, add a little more acid to the buret and allow it to drain through the tip until it is free of bubbles and the meniscus is at 0.0 mL.
4. Repeat step 3 using the base solution in the second buret.

Starting with step 5 of the "Procedure," *one lab partner should carry out the instructions while the second partner records the data.*

5. Place a 125-mL Erlenmeyer flask under the acid buret as in Figure 39-2. Holding a sheet of white paper behind the buret to make the scale easier to read, allow exactly 10.0 mL of acid to flow into the flask.
6. Add exactly 10.0 mL of distilled water to the flask. Then, using a clean dropper pipet, add three drops of phenolphthalein. Swirl the flask to mix all the ingredients.
7. Place the flask on a sheet of white paper under the buret containing the base solution. To avoid splashing, be sure the tip of the buret is in the flask (See Figure 39-2).
8. Swirling the flask gently, begin the titration by adding NaOH to the flask drop by drop. Continue until a faint pink color remains for about 30 seconds. If "overtitration" occurs (the pink color is too deep), follow your teacher's instructions for correcting this condition.
9. Note and record the exact final volume reading on the scale of the base buret. Discard the solution in the flask as instructed. Wash and rinse the flask.
10. Repeat the titration (steps 5 through 9). It is not necessary to refill the burets. Simply read and record the initial volumes of the solutions in the burets carefully.

PART B TITRATION OF WHITE VINEGAR

11. Using a pipet and suction bulb, measure 10 mL of white vinegar into a 250-mL Erlenmeyer flask. Add 100 mL of distilled water.

12. Add three drops of phenolphthalein and *carefully* titrate, using the same NaOH solution used in Part A.

13. If overtitration occurs, add a measured amount of vinegar to the flask (using the pipet) until the solution is colorless. This time, reach the end point carefully by titrating drop by drop with the NaOH solution.

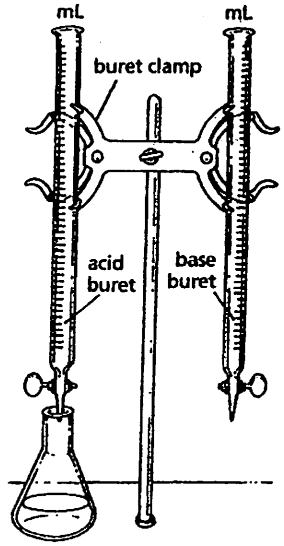

Figure 39-2

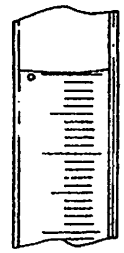
Figure 39-1

Observations and Data

PART A

DATA TABLE

	Trial 1		Trial 2		Trial 3		Trial 4	
	HCl	NaOH	HCl	NaOH	HCl	NaOH	HCl	NaOH
Initial reading								
Final reading								
Volume used								

PART B

Total volume: white vinegar = ____________

Total volume: NaOH solution = ____________

Calculations

PART A

For each trial, calculate the molarity of the NaOH solution using the relationship $M_b = \frac{M_a \times V_a}{V_b}$

Show work on a separate sheet of paper.

Trial 1 ____________

Trial 2 ____________

Trial 3 ____________

Trial 4 ____________

Avg ____________

PART B

To determine the molarity of the NaOH solution used, average the results calculated in Part A. Then use the relationship

$$V_a \times M_a = V_b \times M_b$$

to calculate the molarity of the white vinegar. The volumes of acid and are those in Part B of "Observations and Data."

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Conclusions and Questions

1. How reproducible were the results of your two trials? How did your results compare with those of your partner?

2. Define these terms: standard solution; titration; end point; volumetric analysis; gravimetric analysis.

3. If 30.0 mL of 0.500 *M* KOH is needed to neutralize 10.0 mL of HCl of unknown concentration, what is the molarity of the HCl?

4. How many mL of 0.100 *M* NaOH is needed to titrate 20.0 mL of 0.100 *M* H_2SO_4? Use a balanced equation for the neutralization reaction and explain your calculations.

5. Explain why people can use white vinegar in preparing foods and in cooking without danger to the skin or the internal organs.